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ABSTRACT

This document is an instructional module package prepared in objective form for use by an instructor familiar with general inorganic chemistry. Included are objectives, instructor guides and student handouts. The module considers matter, compounds, chemical equations, solutions, chemical equilibrium, acids and bases and solubility equilibria. (Author/RH)

CHEMISTRY FOR OPERATORS

Training Module 1.320.2.77

Prepared for the

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by

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September, 1977

Module No:

Module Title:

Chemistry for Operators

Approx. Time:

42 hours

Submodule Titles:

Introduction, Compounds, Chemical Equilibria

Solutions, Chemical Equilibrium I, Acids & Bases, Solubility Equilibria

Objectives:

Upon completion of this module the participant should be able to describe the basic structure and types of matter, identify compounds and recognize their various attributes, examine chemical equations and reactions, relate all the various aspects of solutions, explain chemical equilibrium, examine acids and bases and their usage, and explain solubility equilibrium.

Instructional Aids:

Singer SVE filmstrips
Handouts
pH meters
Titration equipments

Instructional Approach:

Lecture
Discussion
Filmstrips
In-class participation

References:

Chemistry Made Simple, by Fred C. Hess.
Chemistry for Laboratory Technicians by Stanley Cherim
Manual of Instruction for Sewage Treatment Plant Operators, by Health Education Service
Chemistry for Sanitary Engineers by Sawyer and McCarty
Instructional Development Workshop, by EPA, Utah State University
Introduction to Chemistry by Turk, Meislich, Brescia, Arents
Analytical Chemistry by Pietryzk and Frank
Modern Chemical Technology by Chemical Technician Curriculum Project
Singer SVE Filmstrip Series (Chemistry) by singer

Class Assignments:

Reading and problem assignments as shown in module

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Introduction
Approx. Time: ½ hour	Topic: Perspectives

Objectives:

Upon completion of this module, the participant should be able to:

1. Define each of the following: Chemistry, chemical properties, chemical change.
2. Describe how chemistry is used in water and wastewater treatment and the significance of chemistry in water and wastewater analysis.

Instructional Aids:

Instructional Approach:

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.

Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.

Class Assignments:

Read PP. 1 - 16, Chemistry Made Simple

Module No:	Topic: Perspectives

Instructor Notes:	Instructor Outline:
pp. 9 - 16 Chemistry Made Simple - Introduction pp. 65 - 69 Chemistry for Laboratory Technicians - Matter	Objective 1 - Definitions a. Chemistry should be presented as a branch of science that deals with the composition of matter and changes in forms of matter. b. Chemical properties are identified as a listing of all the chemical reactions of a substance and should be distinguished from physical properties. c. Chemical change should be described as a change from one substance to another. For example, burning carbon compounds, under oxygen, produces carbon dioxide as one product. This can be used to indicate chemical change and a type of reaction. Emphasize that chemical change involves a change from one substance to an entirely new one and not a physical change, as in water freezing to a solid form. Objective 2 - Significance of Chemistry A sort of overview of chemistry applied to wastewater treatment. Give examples where chemistry may be used such as in coagulation, chlorination, softening, and in laboratory analysis. A lab example could be pH measurements describing that pH actually measures the hydrogen ion concentration in water solutions.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Introduction
Approx. Time: 2 hours	Topic: Types of Matter - Chemical Terms

Objectives: Upon completion of this module, the participant should be able to:

1. Define each of the following: Matter, substances, composition of matter, elements, compounds, mixtures, symbols of elements.
2. Explain that elements are the basic constituents of matter and that all elements known to man are found on the periodic table.
3. Name 10 common elements encountered in water and wastewater and write their symbols.
4. Distinguish between a compound and an element and recognize that compounds are more abundant than elements.

Instructional Aids:

Singer SVE filmstrip - Elements, compounds and mixtures

Singer SVE Filmstrip - The periodic system

Instructional Approach:

Discussion - lecture
Filmstrip - A 493-8
Filmstrip - A 493-2

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Manual of Instruction for Sewage Treatment Plant Operators, Health Education Service
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.

Class Assignments:

Read PP. 16 - 18 Chemistry Made Simple

Module No:	Topic: Types of Matter - Chemical Terms
Instructor Notes:	Instructor Outline:

pp. 16 - 18 Chemistry Made Simple - Chemical Properties

Fig. 12, P. 27 Chemistry Made Simple

P. 93 - 99 Chemistry for Laboratory Technicians Periodic chart

P. 222, Table 11, Manual of Instruction for Sewage Treatment Plant Operators

P. 18 Chemistry Made Simple, Compounds

Objective 1 - Definitions

a. Define matter and explain that types of matter are differentiated in terms of properties.

b. Substances are a definite variety of matter and each specimen has the same properties.

c. Composition of matter relates to the fact that matter is made up of pure substances, elements, or combinations of substances, compounds.

d. Elements are the building blocks of matter, the basic constituents. Define elements by their chemical symbols and give several examples. Compounds should be described as pure substances composed of chemically combined elements. Point out that compounds can be decomposed into elements only by some type of chemical change. Explain mixtures as a continuation of substances held together by physical rather than chemical means.

Objective 2 - Elements, periodic table

The periodic table should be represented as a table of the elements, and identify groups and periods.

Objective 3 - Common Wastewater Elements

Associate elements with their symbols. List about 15 involved in wastewater and discuss where they are encountered.

Objective 4 - Compounds and Elements

Give several examples of compounds by writing their chemical formulas. Compounds and elements should be differentiated by formulas and symbols. Compounds are much more abundant than elements and as elements combine to form compounds, the elements lose their individual properties, and

Module No:	Topic: Types of Matter - Chemical Terms - Cont.
Instructor Notes:	Instructor Outline:

a new set of properties unique to the compound is created. Indicate also that some elements may combine with themselves to form compounds such as chlorine molecules combining to form Cl_2.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Introduction
Approx. Time: 1½ hours	Topic: Structure of Matter

Objectives: Upon completion of this module, the participant should be able to:

1. Differentiate between an element, atom, electron, proton, neutron in terms of electrical charge.
2. Illustrate that atoms of elements have a central "core" called the nucleus consisting of protons and neutrons and electrons orbit the nucleus at relatively large distances in distinct shells.
3. Indicate that electrons in outer shells are involved with chemical reactions and it is these electrons which determine the charge of an atom.
4. Explain that a positive charge balances a negative charge and 2 negative charges balance 2 positive charges etc.
5. Define atomic weight of an element.
6. Given the periodic chart find the symbol for calcium, oxygen, chlorine, flourine, magnesium and their respective atomic weights.

Instructional Aids:

Singer SVE Filmstrip - Group 1, Atoms, molecules, and ions

Instructional Approach:

Discussion
Filmstrip - A 493-1

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.

Class Assignments:

Read PP. 23 - 30 Chemistry Made Simple

Module No:	Topic: Structure of Matter
Instructor Notes:	Instructor Outline:

P. 23 - 30 Chemistry Made Simple, Structure of Matter

P. 82 - 88 Chemistry for Laboratory Technicians - Atomic structure

Objectives 1 and 2 - Bohr Model

Basically describe, by illustration, the Bohr model of the atom and emphasize that electrons, protons, neutrons are the basic particles of an atom: Elements are composed of atoms. Discuss the electroneutraility of elements. Briefly describe atomic number.

Objectives 3 and 4 - Electroneutrality Shells & Charges

Discuss electronic shells of atoms and emphasize that electrons in the outer shell(s) are involved in chemical reactions or changes. Discuss balancing of charges.

Objectives 5 and 6 - Atomic Weight

Atomic weights should be described by using the periodic table. Several examples of chemical names, symbols, and atomic weights should be given. Atomic weights are relative but may be introduced in terms of grams, gram-atomic weight, quantity of an element.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Compounds
Approx. Time: 1½ hours	Topic: Valence

Objectives:

Upon completion of this module, the participant should be able to:

1. Illustrate that the electrons in the outer shells (generally) of the elements can transfer between elements to form compounds.
2. Describe valence and identify the valence number with periodic table groups.
3. Given a list of compounds and ions of these compounds identify the cations and anions.
4. Define radicals and list 8 encountered in water and wastewater.

Instructional Aids:

Handout No. 1

Instructional Approach:

Discussion
Practice finding valence in-class

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.
Manual of Instruction for Sewage Treatment Plant Operators, Health Education Service

Class Assignments:

Read PP. 32 - 37 Chemistry Made Simple
Read Handout

Module No:	Topic: Valence

Instructor Notes:	Instructor Outline:
P. 32 - 39 Chemistry Made Simple, Valence.	Objective 1 - Outer Shell Electrons Describe the process of ion formation in detail. Define it as a transfer of electrons or an electrovalence process. Electrons in outer shells in atoms will need to be re-examined. The participant will not need to work out each ion pair formation, but merely understand the process since valence will generally be memorized or found according to periodic table groups. The sodium chlorine example serves to explain the electron exchange process and the production of ions.
P. 99 - 106 Chemistry for Laboratory Technicians - Nature of the Chemical Bond P. 272 Table 11, Manual of Instruction for Sewage Treatment Plant Operators	Objective 2 - Valence Describe valence number as the number of electrons of the element involved in the formation of a compound. Give the general rules in determining valence number of an element according to periodic table groups. Indicate that valence No. of an ion is numerically equal to the charge on the ion, positive ions have positive valence numbers, negative ions have negative valence numbers. Give several examples. Briefly describe covalence, and differentiate between electrovalence products and covalence products. Define molecules and note that they are present only in covalent compounds.
Handout No. 1	Objective 3 - Cations and Anions Describe cations and anions and give several examples common to wastewater. The importance of ions in solutions could be related in terms of laboratory testing. pH is a measurement of hydrogen ions, hardness measures calcium and magnesium ions, the iron test measures iron ions, and free chlorine measures chlorine ions. Indicate that some compounds when dissolved in solutions, may dissolve into their component cations and anions. List several examples and illustrate how ions are represented.

Module No:	Topic: Valence - Cont.
Instructor Notes:	Instructor Outline:
Handout No. 1	Objective 4 - Radicals Define radicals as chemical compounds or clusters of elements which behave as if they were a single element. List several radicals and their respective valence and indicate that they will always occur as such and must be learned individually through repeated use.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Compounds
Approx. Time: 1 hour	Topic: Nomenclature

Objectives:

Upon completion of this module, the participant should be able to:

1. Name 10 compounds used in water and wastewater treatment and find their formulas.
2. Indicate the use of valence when making compounds and how subscripts are used in formulas to balance charges.
3. Name 5 hardness causing divalent cations and their 5 common anions and make compounds from these.
4. Indicate how radicals are used in formulas.

Instructional Aids:

Instructional Approach:

Discussion
Practice making formulas

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.
Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.

Class Assignments:

Practice making formulas

Module No:	Topic: Nomenclature

Instructor Notes:	Instructor Outline:
P. 35, 36 Chemistry Made Simple, Formulas and Valence P. 109 - 113 Chemistry for Laboratory Technicians - Formulas and Nomenclature Table 12, P. 275 Manual of Instruction for Sewage Treatment Plant Operators - Compound names & formulas Table 19 - 1, P. 349, Sawyer and McCarty, Cations causing hardness	Objectives 2 and 4 - Valence in Formulas, Radicals Indicate that the net sum of all the valence numbers in a given compound must be zero, in most cases. Formulas should be described as the ratio of the atoms of each element present in the compound. Show by several examples how valence charges are used to make formulas for compounds. Show how subscripts are used to balance valences. Use radicals to make formulas and illustrate how they are balanced in respect to charges. The student should be encouraged to practice making formulas and learning valences. The use of suffixes ide, ife, and ate should be briefly discussed with respect to naming compounds. Objectives 1 and 3 - Formulas of Compounds The student should be able to associate formulas of compounds with names. Hardness causing cations and associated anions can be introduced and formulas can be composed for various compounds using these ions.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Compounds
Approx. Time: 2 hours	Topic: Molecular Weights

Objectives:

Upon completion of this module, the participant should be able to:

1. Define molecular weight and indicate how it is computed.
2. Given the actual weight of a compound determine the number of moles.
3. Given 5 compounds commonly found in wastewater select the weight in grams needed to produce 1 molar solutions of these.
4. Indicate that an element may combine with itself to form a molecule.

Instructional Aids:

Handout No. 2
Singer SVE Filmstrip - Atomic and Molecular Weights

Instructional Approach:

Discussion
Handout No. 2
Filmstrip A 493-4

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.
Instructional Development Workshop, EPA, Utah State University, 1973.

Class Assignments:

P. 38 Problem set #4 Chemistry Made Simple
Read Handout No. 2

Module No:	Topic: Molecular Weights
Instructor Notes:	Instructor Outline:

P. 37 - 38 Chemistry Made Simple - Formula or Molecular Weights

Handout No. 2

PP. 114 - 120 Chemistry for Laboratory Technicians, the Molecular concept

Section 5 - Instructional Development Workshop - the molecular concept.

Objective 1 - Molecular Weight

Define formula or molecular weights, use several examples for determination of molecular weights.

Objectives 2, 3, and 4 - MW actual weight relationships

Define molecular quantities and show how the numbers of moles is found given the weight of the compound. Describe molar solutions and show how quantities of compounds are found to make 1 molar solution or fractions of. Allow the participant to work several in class problems. Submodules 1 and 2, Introduction and compounds, should be reviewed at this point and Exam No. 1 may be given at the next session.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Chemical Equations
Approx. Time: 1½ hours	Topic: Equations - Reactions

Objectives:

Upon completion of this module, the participant should be able to:

1. Given a chemical equation define reactants, products.
2. Name and describe four types of chemical reactions.

Instructional Aids:

Singer SVE Filmstrip - Symbols, formulas, and equations

Instructional Approach:

Discussion.
Filmstrip - A 493-3

References:

Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.
Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.

Class Assignments:

Read PP. 20 - 21 Chemistry Made Simple

Module No:	Topic: Equations - Reactions
Instructor Notes:	Instructor Outline:
PP. 69 - 71 Chemistry for Laboratory Technicians, Chemical Change P. 226 Manual of Instruction for Sewage Treatment Plant Operators - Equations P. 20, 21 Chemistry Made Simple Chemical Change	Objective 1 - Chemical Equations Describe the equation form for reactions, reactants, products, symbols used in depicting equations. Briefly introduce the concept of stoichiometry. Give examples of equations commonly used in wastewater treatment. Objective 2 - Types of Reactions Examine the four principal types of chemical reactions which include combination, decomposition, replacement, and double displacement. Discuss energy changes and catalysts involved in reactions. Review how occurrences of chemical change are made obvious such as gas evolution, precipitation, large energy changes, color changes, etc.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Chemical Equations
Approx. Time: 2½ hours	Topic: Balancing Equations

Objectives:

Upon completion of this module, the participant should be able to:

1. Define the law of conservation of matter.
2. Describe how coefficients and subscripts are used in balancing equations to maintain the law of conservation of matter.
3. Given reactants and products from a common reaction write a balanced equation.

Instructional Aids:

Instructional Approach:

Discussion
Practice balancing equations in-class

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.
Instructional Development Workshop, EPA, Utah State University, 1973.
Introduction to Chemistry, Turk, Meislich, Brescia, Arents, 1968.

Class Assignments:

Do problems CH 5, Chemistry Made Simple

Module No:	Topic: Balancing Equations
Instructor Notes:	Instructor Outline:

Chapter 5, Chemistry Made Simple, Laws of Chemistry

PP. 120 - 125, Chemistry for Laboratory Technicians - Equations

Section 10, Instructional Development Workshop, Reading and Writing Chemical Equations-Exercises

Chapter 5, Introduction to Chemistry, Turk - Stoichiometry

Objective 1 - Law of Conservation of Matter

Define the law of conservation of matter and relate it in terms of reactants and products. Also introduce the law of definite properties.

Objective 2 and 3 - Balancing Equations

The rules governing the writing of balanced equations is probably best developed by using several examples. Describe, by examples, how subscripts and coefficients are used to balance equations to maintain the above laws. Examine stoichiometric reactions again as one in which all reactants are converted to final products, irreversible reactions. Introduce reversible reactions and briefly describe the equilibrium concept. Give the student the opportunity to balance several in-class equations when given reactants and products. These objectives will necessarily take some time.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Chemical Equations
Approx. Time: 1 hour	Topic: Ionic Equations

Objectives:

Upon completion of this module, the participant should be able to:

1. Describe the process of ionization in water.
2. Depict with an equation how a compound, when dissolved in water, reacts or ionizes into its component ions.
3. Indicate that the charges on the ions are "balanced" with respect to each other and must be of "equal" and opposite charge.
4. Given several compounds common to wastewater describe by equations how they ionize into their respective ionic components.
5. Indicate that not all compounds ionize to the same degree.

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.

Class Assignments:

Read PP. 76 - 79 Chemistry Made Simple

Module No:	Topic: Ionic Equations

Instructor Notes:	Instructor Outline:
PP. 76 - 78 Chemistry Made Simple - Ionization	Objectives 1, 2, and 3 - Ionization Redefine ions and describe the process of electrolytic dissociation, ionization in solutions. Generally describe the 3 types of electrolytes. Indicate that non-electrolytes do not dissociate into ions but "separate" into molecules in solution. The descriptions of electrolytes and non-electrolytes should be of a general nature dividing electrolytes into acids, bases, and salts which ionize and non-electrolytes which do not ionize. Dissociation of electrolytes may be shown by equations such as: $HCl = H^+ + Cl^-$ $NaCl = Na^+ + Cl^-$ $CaCl_2 = Ca^{++} + 2\ Cl^-$ $Na_2SO_4 = 2\ Na^+ + SO_4 =$ $Al_2(SO_4)_3 = 2\ Al^{+++} + 3\ SO_4 =$ Emphasize that the charge on each ion is the valence number of the atom or radical. Also point out that the charges on the ions are balanced with respect to each other (coefficients and subscripts).
P. 78, 79 Chemistry Made Simple - Strong and Weak Electrolytes	Objectives 4 and 5 - Degree of Ionization Use compounds common to wastewater such as Na_2CO_3, $CaCO_3$, H_2CO_3, which ionize and depict by equations how they form ions in water solutions. Define strong and weak electrolytes. Specify that the terms strong and weak refer to the degree of ionization. Weak electrolytes could be defined as compounds which dissolve in solution only partially as ions and coexist in solution with the molecular forms of the compound. Incomplete ionization may also be described as less than 100 per cent ionization.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solutions
Approx. Time: 3/4 hour	Topic: Introduction

Objectives:

Upon completion of this module, the participant should be able to:

1. Define: Solution, aqueous solution, solute, solvent, mixture.
2. Explain how solutes may be dispersed into molecular, ionic and combinations of the two forms.
3. Differentiate between a solute in solution (homogeneous) and solids suspended in solution (heterogeneous).

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1971.

Class Assignments:

P. 66 Chemistry Made Simple

Module No:	Topic: Introduction
Instructor Notes:	Instructor Outline:
P. 66 Chemistry Made Simple, Solutions P. 162 Chemistry for Laboratory Technicians - Solutions - Introduction	Objective 1 - Definitions Basically objective 1 concerns definitions of terms. Objective 2 - Solutes, molecules and ions. Describe again how a solute coexists as molecules and ionic forms. Objective 3 - Homogeneous and Heterogeneous This concerns differentiating between true solutions and heterogenous solutions such as colloidal systems. Describe colloids.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solutions
Approx. Time: 1½ hours	Topic: Expressing Concentrations

Objectives:

Upon completion of this module, the participant should be able to:

1. Differentiate between strong and weak, concentrated and dilute solutions.
2. Explain dilution factors and give an example using BOD dilutions.
3. Describe molarity as a method for expressing concentration and show how it is calcualted.
4. Prove that concentration unit 1 mg/l = concentration unit 1 ppm.

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1972.
Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.

Class Assignments:

Work concentration problem (Objective 4) as assigned
Read PP. 66 - 67 Chemistry Made Simple
Work practical dilution problems

Module No:	Topic: Expressing Concentrations
Instructor Notes:	Instructor Outline:

PP. 166 - 174 Chemistry for Laboratory Technicians - Concentrations of solutions

P. 66, 67 Chemistry Made Simple, Methods of Expressing Concentration - Dilution

P. 403 Sawyer and McCarty, BOD dilutions

Objectives 1 and 2 - Dilutions

Explain concentrations in terms of concentrated dilute, and percent solute. Examine how practiced dilution problems may be solved; for example diluting concentrated stock solutions or samples for laboratory analysis. BOD dilutions could be used to examine how dilution factors are determined.

Objective 3 and 4 - Molarity, concentrations

Redefine the relationship between molarity, mole, and volume. Show mathematically how mg/l is converted into parts per million. Give molar concentrations of solute, then convert to mg/l and ppm.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solutions
Approx. Time: 3 hours	Topic: Normality

Objectives:
Upon completion of this module, the participant should be able to:

1. Define normality as a method of expressing concentration.
2. Given the normality of a compound describe and compute the number of equivalents and show the relationship between equivalents, actual weight of compound, and equivalent weight.
3. Indicate a fundamental law of chemistry which states that a given number of equivalents will react exactly with the same number of equivalents of another substance.
4. Differentiate between differing normalities in terms of concentration, 1N, 6 N, 12 N.

Instructional Aids:

Handout No. 3

Instructional Approach:

Discussion
Handout No. 3 - Work problems

References:

Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1972.
Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Manual of Instruction for Sewage Treatment Plant Operators, Health Education Service
Analytical Chemistry, Pietryzk and Frank, Academic Press, 1974.

Class Assignments:

Read PP. 175 - 180 Chemistry for Laboratory Technicians
Study Handout

Module No:	Topic: Normality

Instructor Notes:	Instructor Outline:
PP. 175 - 180 Chemistry for Laboratory Technicians - Normal Solutions P. 66 Chemistry Made Simple - Normality P. 229 - 230 Manual of Instruction for Sewage Treatment Plant Operators, Normal Solution Handout No. 3 PP. 56 - 59 Analytical Chemistry - Calculations based on Normality	Objective 1 and 2 - Normality Review valence numbers for several examples. Indicate that in a stoichometric reaction a constituent whose amount is being measured undergoes a reaction with another substance or is made to decompose in accordance with a well defined equation written in terms of reactants and products. Review equation forms, generally $R_A + R_B \longrightarrow R_C + R_D$. Describe that normality, like molarity, is a method of expressing concentration and is commonly used in wastewater control. Define equivalent weight and emphasize that it is dependent upon the valence involved in the reaction. Use examples and refer to Handout No. 3. Determine equivalents from equivalent weight, and then describe normality. Use handout examples. Show that normality then is an alternate method, though an indirect one, of expressing concentration. Its utility is defined by the fundamental law $N_1 \times ml_1 = N_2 \times ml_2$, described in (2). Work through the relationships of N, equivalents, equivalent weight, and actual weight several times. Objective 3 - Equivalents in reactions The significance of using normality is useful due to the fact that one equivalent of reactant A will react with exactly 1 equivalent of reactant B, 4 equivalents with 4 equivalents, and so on. Present the fundamental equation describing this relationship. Normality 1 x volume 1 = Normality 2 x Volume 2, equivalents = equivalents. Use an example such as what is the normality of a Na OH solution if a specified ml is needed to exactly react with a specified g quantity of an acid, HCl?

Module No:	Topic: Normality
Instructor Notes:	Instructor Outline:

Objective 4 - Concentration in terms of normality

Show how differing normalities of a substance describes differing concentrations and discuss their preparation.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solutions
Approx. Time: 1½ hours	Topic: Solubility

Objectives:
Upon completion of this module, the participant should be able to:

1. Define solubility and factors influencing solubility in terms of concentration.
2. Explain the process of dissolving and deposition of a solid in solution.
3. Distinguish between a saturated, unsaturated and super-saturated solution.
4. Explain the equilibrium condition for substances at saturation where dissolving and deposition processes occur at equal rates simultaneously.

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.
Chemistry Made Simple, Fred C. Hess, Doubleday and Company, 1955.
Instructional Development Workshop, EPA, Utah State University.
Introduction to Chemistry, Turk, Meislich, Brescia, Arents, 1968.

Class Assignments:

Read PP. 69, 70 Chemistry Made Simple
Do Section 18 Instructional Development Workshop

Module No:	Topic: Solubility
Instructor Notes:	Instructor Outline:

P. 69, 70 Chemistry Made Simple - Solubility

Section 18 - Instructional Development Workshop - The Nature of Solutions

P. 31 Sawyer and McCarty, Heterogeneous Chemical Equilibria

Objective 1 - Solubility

Describe solubility and factors influencing solubility. I would restrict the factors influencing solubility to solids in liquids. Explain solubility in terms of concentration, mg/l.

Objective 2 and 3 - Saturated, Unsaturated Solutions

Illustrate the reversible process of dissolving and deposition of a solid in solution. Explain the conditions of saturated, unsaturated, and supersaturated for solids in liquids.

Objective 4 - Solubility Equilibria

Describe heterogeneous equilibria for solids and indicate relative saturation values. Specify that the dissolved portion is ionic. Show by equation from the reversible relationship between ions in solution and the deposited solid and relate to an illustration.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solutions
Approx. Time: 3/4 hour	Topic: Ionization in Solutions

Objectives:

Upon completion of this module, the participant should be able to:

1. Describe the ionization process in aqueous solution.
2. Explain how some solutions will conduct electricity and others will not and how this is used in laboratory testing.

Instructional Aids:

Conductivity demonstration if available

Instructional Approach:

Discussion
Demonstration - conductivity

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Co., 1955.
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1972.

Class Assignments:

Module No:	Topic: Ionization in Solutions
Instructor Notes:	Instructor Outline:

P. 77 Chemistry Made Simple - Ionization

P. 239 Chemistry for Laboratory Technicians - Conductimetry

Objective 1 - Ionization

Re-examine the ionization process in water solutions. Define degree of ionization and indicate that much of laboratory testing is concerned with ions in solutions.

Objective 2 - Conductivity - Solutions

Explain the analysis of total ionic concentrations by conductivity measurements. Specify that the purity of water may be determined in part by this measurement.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solutions
Approx. Time: 1 hour	Topic: Suspensions, Colloids, Solids

Objectives:

1. Differentiate between true solutions and colloidal dispersions and name types of substances in water that are sometimes considered soluble actually form colloidal dispersions.
2. Define turbidity and describe causes.
3. Differentiate between dissolved substances and undissolved substances.
4. Differentiate between total solids and suspended solids in water and wastewater as performed by lab analysis.

Instructional Aids:

Solids testing apparatus
Singer SVE Filmstrip - solutions, suspensions, and colloids

Instructional Approach:

Discussion
Filmstrip - A 493-6

References:

Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.

Class Assignments:

Read PP. 290 - 298, 214, 215 Chemistry for Sanitary Engineers.

Module No:	Topic: Suspensions, Colloids, Solids

Instructor Notes:	Instructor Outline:
P. 214, 215 Sawyer and McCarty Colloid Chemistry	Objectives 1 and 3 - True and Colloidal Describe the relationship between colloidal dispersions and true solutions, molecular and ionic. Use colloidal dispersions to exemplify the narrow range of sizes between dissolved and undissolved substances.
PP. 290 - 298 Sawyer and McCarty - Solids	Objective 2 - Turbidity Define turbidity and indicate the range of sizes of suspended materials causing turbidity.
PP. 435 - 437 Sawyer and McCarty - Solids	Objective 4 - Solids Use solids analysis to diagrammatically show the relationship of total, suspended and dissolved solids.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solutions
Approx. Time: 3/4 hour	Topic: Miscellaneous Solubility

Objectives:

Upon completion of this module, the participant should be able to:

1. Using dissolved oxygen and carbon dioxide describe the relationship between pressure and temperature and the solubility of these gases in a lagoon.
2. Contrast between the relative solubilities of heavy metals, pesticides, inorganic and organic substances in water.

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Co., 1955.

Class Assignments:

Read PP. 69 - 71 Chemistry Made Simple

Module No:	Topic: Miscellaneous Solubility
Instructor Notes:	Instructor Outline:

P. 69 - 71 Chemistry Made Simple - Solubility

Objective 1 - Solubility of Gases

Explain the factors influencing solubility of gases in liquids. Using lagoons show the relationships of these factors to solubility of dissolved oxygen and carbon dioxide.

Objective 2 - Relative Solubilities of Substances

Discuss the fact that solubility of substances in water is directly related to the nature of the solute. Discuss the relative solubility of substances such as pesticides, relatively insoluble, metals fairly soluble, organics generally less soluble than inorganics etc.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solutions
Approx. Time: 3/4 hour	Topic: Miscellaneous Solubility.

Objectives:

Upon completion of this module, the participant should be able to:

1. Using activated carbon and other materials describe the process of adsorption of solutes.
2. Indicate the employment of osmosis and dialysis in water and wastewater treatment.
3. Describe miscibility, partitioning of solvents, and solvent extractions.

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.
Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1972.
Chemistry Made Simple, Fred C. Hess, Doubleday and Co., 1955.

Class Assignments:

Read PP. 206 - 209, 177 - 180 Chemistry for Sanitary Engineers
Read P. 71 Chemistry Made Simple

Module No:	Topic: Miscellaneous

Instructor Notes:	Instructor Outline:
PP. 206 - 209 Sawyer and McCarty, Adsorption	Objective 1 - Adsorption Define adsorption and explain why activated carbon is used extensively; large surface area in relation to mass.
PP. 177 - 180 Sawyer and McCarty, Osmosis and Dialysis	Objective 2 - Osmosis and Dialysis Briefly describe osmosis and indicate how dialysis is used in separation of solutes.
P. 163 - 166 Chemistry for Laboratory Technicians - Solubility and Miscibility P. 71 Chemistry Made Simple - Law of Partition	Objective 3 - Solvent Miscibility Briefly describe miscibility, solvent partitioning, and extractions. Exam No. 2 - Next Session

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Chemical Equilibrium I
Approx. Time: 2½ hours	Topic: General Equilibrium Conditions

Objectives:

Upon completion of this module, the participant should be able to:

1. Define reversible and non-reversible reactions.
2. Explain the reversible reaction condition for solids in a saturated solution.
3. Given a reaction describe the equilibrium equation and set up a mathematical expression for the equilibrium constant.
4. State LeChatelier's principle and indicate how it is used to predict the outcome of a given change in equilibrium conditions.
5. Differentiate between heterogeneous and homogeneous equilibrium.

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Co., 1955.
Instructional Development Workshop, EPA, Utah State University.

Class Assignments:

Read Section 6 Instructional Development Workshop
Read PP. 78 - 84 Chemistry Made Simple

Module No:	Topic: General Equilibrium
Instructor Notes:	Instructor Outline:

PP. 78 - 84 Chemistry Made Simple, Chemical Equilibria

Section-6, Instructional Development Workshop - Chemical Equilibria

Objectives 1 and 3 - Reversible, Non-reversible Reactions, Equilibrium Expression

Describe the condition of equilibrium for partially ionized weak electrolytes as an example of reversible reactions. Distinguish equilibrium types from non-reversible reactions. Use equations and examine the symbols used, molar concentrations, equilibrium expression, etc.

Acetic acid could be used to show, by equation, how it dissociates partially into ions resulting in molecular-ionic equilibrium.

Differentiate again between strong and weak electrolytes, percent ionization.

Explain how a weak electrolyte maintains a constant degree of ionization, and set up the mathematical expression. Show the relationship between initial compound concentration and resultant ionic concentrations using the ionization constant equation.

Show how coefficients in the reaction equation are used in the ionization constant equation. Work out equilibrium expressions for situations where coefficients are used or not used. The ionization equation may be defined as the law of mass action expression.

Objective 2 - Solids - Reversible Reaction

Explain the reversible reaction situation for a slightly soluble solid. Discuss the equilibrium condition of solids in equilibrium with their ions by illustrating with a slightly soluble solid. This is covered in more detail in topic Solubility Equilibria.

Objective 4 - LeChatelier's Principle

With emphasis on concentration, show how an equilibrium situation for a weak electrolyte will shift position of equilibrium to adjust

Module No: -	Topic: General Equilibrium
Instructor Notes:	Instructor Outline:

for a change in concentration. For example, using HF, describe how an increase in F ions from an external source will introduce stress and cause the position of equilibrium to shift in relieving the stress. Mathematical treatment is probably not necessary.

Objective 5 - Heterogeneous, homogeneous Equilibrium.

Briefly discuss the differences between heterogeneous and homogeneous equilibrium.

Module No: 2	Module Title: Chemistry for Operators
	Submodule Title: Acids and Bases
Approx. Time: 2 hours	Topic: Introduction

Objectives:

Upon completion of this module, the participant should be able to:

1. Define acid, base, salt, conjugate acid and base, neutralization, hydronium ion
2. List common acids and bases used in water and wastewater.
3. When given the name or formula for an acid and base write a balanced chemical equation for the reaction which occurs and the name the salt formed.
4. Differentiate between strong and weak acids and bases.

Instructional Aids:

Singer SVE Filmstrip - acids, bases, and salts

Instructional Approach:

Discussion
Filmstrip - A 493-7

References:

Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1972.
Instructional Development Workshop, EPA, Utah State University.
Manual of Instruction for Sewage Treatment Plant Operators, Health Education Service.

Class Assignments:

Read PP. 186 - 188 Chemistry for Laboratory Technicians.

Module No:	Topic: Introduction

Instructor Notes:	Instructor Outline:
P. 186 - 188 Chemistry for Laboratory Technicians - Acids and Bases P. 227, 228 Manual of Instruction for Sewage Treatment Plant Operators - Acids & Bases Section 20 Instructional Development Workshop - Acids and Bases	Objective 1 - Acids, bases Define acids and bases as proton donors and acceptors. Define neutralization as a term applying to equal numbers of H and OH ions combining to form water. Objective 2 - Common acids and bases List acids and bases common to wastewater treatment or analysis such as sulfuric, carbonic, hydrochloric, sodium hydroxide, ammonium hydroxide etc. Objective 3 - Neutralization Show neutralization reactions by equations for simple acid - base titrations. Objective 4 - Strength of acids and bases Emphasize that strong and weak acids and bases refers to the degree of ionization.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Acids and Bases
Approx. Time: 1½ hours	Topic: pH

Objectives:

Upon completion of this module, the participant should be able to:

1. Define pH and indicate that pH is a mathematical expression for hydrogen ion concentration.
2. Illustrate the pH scale diagrammatically showing neutral, increasing basic strength and increasing acid strength.
3. Identify 3 instances in water and wastewater treatment where pH measurements are used and indicate their significance.

Instructional Aids:

pH meter, pH colorimetric testing kit

Instructional Approach:

Discussion
Demonstration - pH meter, colormetric pH tests

References:

Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1972.
Instructional Development Workshop, EPA, Utah State University.
Chemistry Made Simple, Fred C. Hess, Doubleday and Co., 1955.
Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.
Modern Chemical Technology, Chemical Technician Curriculum Project, Volume 3, 1971.

Class Assignments:

Read PP. 84, 85 Chemistry Made Simple
Read CH 15, Chemistry for Sanitary Engineers

Module No:	Topic: pH

Instructor Notes:	Instructor Outline:
PP. 192-196 Chemistry for Laboratory Technicians - the concept of pH Section B Instructional Development Workshop, pH Demonstration with pH meter, colorimetric tests P. 84, 85 Chemistry Made Simple pH PP. 437 - 447 Modern Chemical Technology - Measurement of pH Ch. 11 Sawyer and McCarty, pH	Objectives 1 and 2 - pH Describe pH as a function of hydrogen ion concentration. Give the mathematical expression but limit mathematical treatment. Explain how pH may be measured by using pH meter or colorimetric methods. Illustrate the pH scale. Give examples of substances as they relate to a certain pH. Objective 3 - pH Measurements - Wastewater Simply discuss examples of where pH is taken in wastewater treatment, such as in a digestor, lagoon, effluents, etc. and briefly explain why they are taken.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Acids and Bases -
Approx. Time: 1 3/4 hours	Topic: Neutralization Titrations

Objectives:

Upon completion of this module, the participant should be able to:

1. Define neutralization reactions, titration, end point, indicators.
2. Differentiate between acidity, alkalinity and pH.
3. Discuss how acidity or alkalinity can be determined using the basic equation $N_1 \times V_1 = N_2 \times V_2$.
4. Describe how alkalinity in terms of Mg/liter as calcium carbonate can be determined by titration with a specified normality of acid titrant.

Instructional Aids:

Equipment for titration - acidity

Instructional Approach:

Discussion
Titration demonstration

References:

Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1972.
Manual of Instruction for Sewage Treatment Plant Operators, Health Education Service.

Class Assignments:

Read PP. 189 - 192 Chemistry for Laboratory Technicians

Module No:	Topic: Neutralization

Instructor Notes:	Instructor Outline:
PP. 189 - 197 Chemistry for Laboratory Technicians - Acid base titrations Demonstration of a titration	Objective 1 and 3 - Determining acidity and alkalinity Simply discuss the mechanics of neutralization titrations, how endpoints are determined. Using the law of mass action $N_1 \times V_1 = N_2 \times V_2$ show how acidity is determined. Give a basic equation form whereby so many milliliters of titrant will yield a quantity of acid. Demonstrate titration of a sample for total acidity. Objective 2 - pH, acidity, alkalinity Emphasize that pH is not necessarily a measurement of acidity or alkalinity but merely measures the activity of the hydrogen ion; and not necessarily the total concentration of an acid or base since ionization is a factor. Objective 4 - Determining alkalinity - Explain that acidity is reported in terms of equivalent of Ca CO_3 since many times it is not known what acids are present. Show how acidity mg/l is changed to mg/l Ca CO_3.
P. 230 Manual of Instruction for Sewage Treatment Plant Operators, acidity, alkalinity, pH	

Module No: —	Module Title: Chemistry for Operators
	Submodule Title: Acids and Bases
Approx. Time: 1 hour	Topic: Incomplete Ionization

Objectives:

Upon completion of this module, the participant should be able to:

1. Indicate that weak acids and bases are incompletely ionized and coexist in solution as molecules and ions.
2. List 2 strong and weak acids and bases found in water and wastewater.
3. Given the ionic equation, describe the equilibrium condition that occurs in a weak acid or base in solution and set up a mathematical expression for the ionization constant.
4. Illustrate how a salt of an acid or base contains a common ion.

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry Made Simple, Fred C. Hess, Doubleday and Co., 1955.

Class Assignments:

Read PP. 86, 87 Chemistry Made Simple

Module No:	Topic: Incomplete Ionization.
Instructor Notes:	Instructor Outline:
P. 86, 87 Chemistry Made Simple - Hydrolysis	Objective 1 - Incomplete ionization This concerns reiteration of previous applications. Objective 2 - Acids and bases, examples Give examples of strong and weak acids and bases. Objective 3 - Equilibrium expression Review equilibrium expressions for weak electrolytes and the ionization constant expression. Objective 4 - Common ions Show by neutralization reactions how salts carry an ion common to the acid and base.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Acids and Bases
Approx. Time: 1 3/4 hours	Topic: Buffers

Objectives:

Upon completion of this module, the participant should be able to:

1. Define pH buffers.
2. Explain how buffers are used to resist changes in pH.
3. Describe how buffers are prepared.
4. Given the addition of an acid or bases to a buffer solution describe the mechanism involved in determining the capacity to buffer.

Instructional Aids:

pH meter, buffer solutions, standard acid solutions
Film if available

Instructional Approach:

Discussion
Demonstration of buffers action - titration

References:

Chemistry for Laboratory Technicians, Stanley Cherim, W. B. Saunders Co., 1972.
Instructional Development Workshop, EPA, Utah State University.
Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.
Modern Chemical Technology, Chemical Technician Curriculum Project.

Class Assignments:

Read PP. 196 - 198 Chemistry for Laboratory Technicians

Module No:	Topic: Buffers

Instructor Notes:	Instructor Outline:
PP. 196 - 198 Chemistry for Laboratory Technicians - Buffers and pH PP. 70 - 73 Sawyer and McCarty Buffers Section 10 Instructional Development Workshop, the common ion effect and buffered solutions PP. 447 - 449 Modern Chemical Technology - Buffers	Objectives 1 and 2 - pH buffers Define buffers and how they act to resist changes in pH. Indicate the importance of natural buffer systems in wastewater. Specify the common ions in buffer systems. Objectives 3 and 4 - Mechanism of buffering Describe how buffers are prepared and show the actual process of buffering involved for a buffer solution. Explain why buffers are used to standardize pH meters. A buffer demonstration could be presented by preparing a buffer solution and adding acid. Examine pH changes by a pH meter. Compare this with a non-buffered solution by adding acid to it and observe pH changes.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Acids and Bases
Approx. Time: 1½ hours	Topic: Alkalinity in Natural Waters

Objectives:

Upon completion of this module, the participant should be able to:

1. List contributors of alkalinity in natural waters.
2. Describe why alkalinity as a measure of buffer capacity is so important in wastewater treatment.
3. Explain how algae blooms cause high pH values.
4. Describe how P alkalinity and total alkalinity is determined by titration with sulfuric acid.

Instructional Aids:

Titration equipment

Instructional Approach:

Discussion
Alkalinity titration demonstration

References:

Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.

Class Assignments:

Module No:	Topic: Alkalinity - Natural Waters
Instructor Notes:	Instructor Outline:
PP. 327 - 340 Sawyer and McCarty Alkalinity.	Objective 1 and 2 - Alkalinity contributors Define alkalinity as a measure of its capacity to neutralize acids. Natural waters' alkalinity is primarily due to salts of weak acids. Describe the forms of alkalinity in natural waters such as bicarbonates, salts of phosphates etc.
	Objective 3 - Algae blooms and pH Show how algae blooms in lagoons may increase the pH value, describe changes in alkalinity forms by equilibrium equations for this situation.
Demonstration of titration	Objective 4 - Titration for alkalinity Demonstrate how alkalinity is determined by titration and equate to a simple formula, ml titrant vs. a quantity expressed as mg/l alkalinity.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Acids and Bases
Approx. Time: 3/4 hour	Topic: Acidity in Natural Waters

Objectives:

Upon completion of this module, the participant should be able to:

1. Name the most common sources of acidity in natural waters.
2. Discuss why ground waters contain large amounts of CO_2.

Instructional Aids:

Instructional Approach:

Discussion

References:

Chemistry for Sanitary Engineers, Sawyer and McCarty, McGraw Hill Co., 1967.

Class Assignments:

Read PP. 321, 322 Chemistry for Sanitary Engineers

Module No:	Topic: Acidity - Natural Waters
Instructor Notes:	Instructor Outline:
PP. 321, 322 Sawyer and McCarty Sources and Nature of Acidity	Objective 1 - Acidity sources Discuss sources of acidity in natural waters. Objective 2 - CO_2 in ground water Indicate the reasons for high concentrations of CO_2 in ground waters.

Module No:	Module Title: Chemistry for Operators
	Submodule Title: Solubility Equilibrium
Approx. Time: 2½ hours	Topic: Solubility Equilibria

Objectives:

Upon completion of this module, the participant should be able to:

1. Explain that insoluble is "relative" and means slightly soluble since all salts dissolve to some extent.
2. Given calcium carbonate as an example, explain that as "insolubles" dissolve they do so as ions and result in solid-ionic equilibrium.
3. Given an insoluble compound describe the equilibrium equation and the mathematical expression for the solubility constant.
4. Discuss how the addition of a common ion will upset the equilibrium and how the stress is relieved.
5. Examine how slightly soluble compounds have differing degrees of solubility.

Instructional Aids:

Instructional Approach:

Discussion

References:

Introduction to Chemistry, Turk, Meislich, Brescia, Arents, 1968.
Chemistry Made Simple, Fred C. Hess, Doubleday and Co., 1955.
Analytical Chemistry, Pietryzk and Frank, Academic Press, 1974.

Class Assignments:

Read PP. 87 - 89 Problems Set # 11, Chemistry Made Simple

Module No:	Topic: Solubility Equilibria

Instructor Notes:	Instructor Outline:
Section 7, Instructional Development Workshop - Solubility Equilibria PP. 87 - 89 Chemistry Made Simple - Solubility Product	Objective 1, 2, and 3 - Insoluble solids, solubility equilibria Redefine solubility and insoluble solids, point out that "insoluble" is a relative term since all salts dissolve to an extent. Re-examine the reversible situation for solid-ionic equilibrium. Define the mathematical expression for solubility products and define symbols used. Objective 4 - Solubility, common ion Show how the addition of a common ion from an external source will upset the equilibrium condition and result in deposition of the solid. Objective 5 - Degrees of solubility Discuss relative differences between solubilities and solubility products for unlike compounds. Indicate that some compounds, which contain a common ion, may precipitate out sooner or later than others depending upon the magnitude of the solubility product. Exam No. 3

Module No:	Module Title:
	Chemistry for Operators
	Submodule Title:
Approx. Time:	Introduction and Compounds
1½ hours	EVALUATION - Part A - Covering Submodules Introduction and Compounds

Objectives:

Upon completion of this module the participant should be able to answer correctly 70% of the evaluation questions.

Evaluation Questions:

1. Using the periodic table find the symbols for the following elements and list their atomic weights: Calcium, sodium, nitrogen, iron, chlorine, magnesium, mercury.

2. Given the compound $CaCO_3$ list the elements it is composed of and the relative quantities of atoms in the compound.

3. Describe the general positions of neutrons, protons, and electrons of atoms in elements, relative to each other. Differentiate in terms of electrical charge.

4. Which electrons of the shells of an atom are involved with chemical reactions.

5. For the compounds $CaCO_3$, Na Cl, Na_2SO_4, HCl, $Ca(OH)_2$, identify cations and anions which these compounds may produce. List the valences.

6. Define radicals and give the valences of the following radicals: OH, SO_4, NO_3, CO_3.

7. Make formulas for the following:

   Na and SO_4 Valences Na = +1 SO_4 = -2
   Ca and Cl Valences Ca = +2 Cl = -1
   Ca and OH Valences OH = -1
   Ca and HCO_3 Valence HCO_3 = -1
   Na and CO_3 Valence CO_3 = -2

   Note that subscripts are necessary

8. Which of the following cations are generally associated with water hardness:

   Ca^{++}, Mg^{++}, Fe^{++}, Hg^{++}, Pb^{++}, Ba^{++}

9. Find the molecular weight of the following:

   $CaCO_3$, $Ca(HCO_3)_2$, Na OH, H_2SO_4, SiO_3, Cl_2

10. Given 20 grams $CaCO_3$, 10 grams Na OH, 33 g. Ca $(HCO_3)_2$ compute the number of moles for the compounds.

11. Define a 1 molar solution of Na OH, .1 molar, .01 molar.

12. State whether the following is a compound or element:

Cl _____
Cl_2 _____
Fe _____
Na OH _____
Ca $(OH)_2$ _____
F _____
Sulfuric acid _____
Calcium bicarbonate _____

13. Indicate whether the following statement is true or false. Water is composed of 2 atoms of hydrogen and 1 atom of oxygen.

Module No:	Module Title:
—	Chemistry for Operators
Approx. Time:	Submodule Title: Chemical Equations and Solutions
1¼ hours	EVALUATION - Part B - Covering Chemical Equations and Solutions

Objectives:

Upon completion of this module the participant should be able to answer correctly 70% of the evaluation questions:

Evaluation Questions.

1. List the reactants and products of the following reaction: HCl + Na OH Na Cl + H_2O.
2. Show by equation form how hydrogen flouride ionizes into its component ions.
3. True or False. All compounds that ionize do so to the same extent.
4. Indicate whether true or false.

   a. T F Electrolytes dissociate into ions in solution.
   b. T F Non-electrolytes dissociate into molecules in solution.

5. Given a solution of Na OH (sodium hydroxide) in water which is the solute and solvent.
6. Define aqueous solution.
7. Define the difference between concentrated and dilute solutions in terms of amount of solute in solution.
8. True or False. A .1 molar solution is more concentrated than a 1 molar solution.
9. A .25 molar solution of H_2SO_4 contains how many moles in 1 liter, grams in 1 liter?
10. Given a .3 normal solution of Na OH how many equivalents per liter does it contain.
11. Which of the following normalities of H_2SO_4 (sulfuric acid) is more concentrated, 1N, 6N 12N?
12. Can solutes be dispersed in water solutions into molecular and ionic forms at the same time? Yes or No.
13. Find the equivalent weight of the following: $CaCO_3$, Na OH, HCl.
14. Given 10 g. of $CaCO_3$ dissolved in 1 liter compute the number of equivalents of $CaCO_3$ and its normality.

15. Define solubility for solids in liquids and distinguish between saturated, unsaturated, and supersaturated in terms of relative quantities of solute in solvent.

16. Total ionic content can be measured by what method?

17. Total solids minus suspended solids equal ideally what form of solids.

18. True or false. As the temperature of a lagoon increases the solubility of dissolved oxygen decreases.

Module No:	Module Title:
1	Chemistry for Operators
	Submodule Title:
Approx. Time:	Chemical Equilibrium, Acids & Bases, Solubility Equilibrium
1 hour	EVALUATION - Part C - Covering acids & bases, solubility equilibria

Objectives:

Upon completion of this module the participant should be able to answer correctly 70% of the evaluation questions.

1. Define reversible and non-reversible reactions.
2. Show the mathematical expression for the equilibrium constant for the weak acid HF.
3. Explain what an increase in H^+ ions in No. 2 above will do to the equilibrium situation.
4. Define acid and base in terms of protons.
5. Give the chemical equation for the reaction between Na OH and HCl, show products.
6. a. T or F. Strong acids ionize 100%.
   b. T or F. Weak acids ionize 100%.
7. Briefly state LeChatelier's principle.
8. pH is a function of _____concentration.
9. Illustrate the pH scale showing neutral, basic and acidic areas of pH.
10. Given 10 ml of acid HCl (.1N) how many ml of .01 N Na OH are required to react to neutralize the acid. Use the equation $N_1 \times V_1 = N_2 \times V_2$.
11. List whether the following are weak or strong acids. H_2SO_4; H_2CO_3, HCl; HNO_3
12. Define pH buffers.
13. Name the common ion which HCl and Na Cl contains.
14. Briefly explain why ground waters contain relatively large amounts of CO_2.
15. Explain why algae blooms may cause high pH values.
16. True or False. All solids are soluble to a certain degree, some more than others.
17. Show the mathematical expression for the equilibrium constant for the slightly soluble salt Ba SO_4 (Barium sulfate), and define K_{sp}. Ba^{++} ions and SO_4 = ions.

18. For a saturated solution of Ba SO_4 how will an increase in SO_4 ions from an external source upset the equilibrium situation and how is the "stress" relieved.

19. Activated carbon is sometimes used to _____ (adsorb, or absorb) solutes out of solution.

20. Define osmosis and dialysis.

Module No:	Module Title: Chemistry for Operators
Approx. Time: 1½ hours	Submodule Title: Introduction and Compounds
	EVALUATION - Part A - Covering Submodules Introduction and Compounds

Objectives:

ANSWERS

1. Ca - 40.08
   Na - 22.9898
   N - 14.0067
   Fe - 55.847
   Cl - 35.453
   Mg - 24.305
   Hg - 200.59

2. Ca CO_3 is composed of Ca (calcium), C (carbon) and O (oxygen). Relative quantities are 1 atom Ca, 1 atom C, 3 atoms O.

3. Elements are composed of atoms. Atoms consist of a nucleus core which contains neutrons and protons and clouds or shells of electrons orbiting the nucleus. The atomic number determines the quantities of neutrons, protons, and electrons of an atom. Protons - positive charge, electrons - negative charge, neutrons - neutral (no charge).

4. Electrons in the outer shells of an atom.

5.

	cations	(valence)	anions	(valence)
$CaCO_3$	Ca	(+2)	CO_3	(-2) radical
NaCl	Na	(+1)	Cl	(-1)
Na_2SO_4	Na	(+1)	SO_4	(-2) radical
HCl	H	(+1)	Cl	(-1)
$Ca(OH)_2$	Ca	(+2)	OH	(-1) radical

6. Radicals consist as groups of elements which will always act as a single unit in a chemical compound or reaction.

   OH (hydroxide) -1

   SO_4 (sulfate) -2

   NO_3 (nitrate) -1

   CO_3 (carbonate) -2

7. Na_2SO_4 Sodium sulfate

   $Ca\ Cl_2$ calcium chloride

   $Ca\ (OH)_2$ calcium hydroxide

   $Ca\ (HCO_3)_2$ calcium bicarbonate

   Na_2CO_3 sodium carbonate

8. Ca^{++}, Mg^{++}, Fe^{++}

9. $CaCO_3$ = 100.0892 (100)

   $Ca\ (HCO_3)_2$ = 223.1313 (223)

   NaOH = 39.9971 (40)

   H_2SO_4 = 98.0734 (98)

   SiO_3 = 76.0842 (76)

   Cl_2 = 70.906 (71)

10. .2 moles $CaCO_3$

    .25 moles NaOH

    .148 moles $Ca\ (HCO_3)_2$

11. 1 molar solution of NaOH is 1 mole of NaOH per liter of solvent, .1 molar is .1 mole NaOH per liter solvent, .01 molar is .01 mole NaOH per liter solvent.

12. Cl - element

    Cl_2 - compound

    Fe - element

    NaOH - compound

    $Ca\ (OH)_2$ - compound

    F - element

    Sulfuric acid - compound

    Calcium-bicarbonate - compound

13. Water - H_2O, True

Module No:	Module Title:
—	Chemistry for Operators
	Submodule Title:
Approx. Time:	Chemical Equations and Solutions
1¼ hours	EVALUATION - Part B - Covering Chemical Equations and Solutions

Objectives:

ANSWERS

1. HCl, NaOH - reactants

   NaCl, H_2O - products

2. Hf $H^+ + F^-$

3. False

4. a. True
   b. True

5. NaOH = solute

   H_2O = solvent

6. Aqueous solution is a solution of solutes in a water solvent. Aqueous is derived from water.

7. Concentrated contains more solute in solution than dilute solutions. A concentrated solution contains a substantial quantity of solute, a dilute solution contains a relatively small quantity of solute relative to the quantity of solvent.

8. False

9. .25 moles, 24.5 grams

10. .3 equivalents

11. 12 N

12. Yes

13. $CaCO_3$ equivalent weight = 50.

    NaOH equivalent weight = 40

    NCl equivalent weight = 36.5

14. .2 equivalents, .2 N

15. Solubility of a solid in a liquid is the quantity of solid that dissolves in a given quantity of solvent to form a saturated solution. Saturated refers to the "normally" maximum amount of solute allowed to dissolve in a specified quantity of solvent, solubility. Unsaturated refers to a less than attainable quantity of solute per quantity of solvent, and supersaturated refers to an over maximum quantity of solute per given quantity of solvent.

16. Conductivity, conductometric measurements.

17. Dissolved solids

18. True

Module No:	Module Title:
	Chemistry for Operators
Approx. Time:	Submodule Title:
	Chemical Equilibrium, Acids & Bases, Solubility Equilibrium
1 hour	EVALUATION - Part C - Covering acids & bases, solubility equilibria

Objectives:

ANSWERS

1. Reversible reactions are those which produce products from reactants and may also reverse to produce reactants from products. An equilibrium condition is attained whereby both situations occur simultaneously. Non-reversible reactions are those which go to completion, all reactants are used up "for the reaction" to produce products and there is no equilibrium situation.

2. $HF \quad H^+ + F^-$

$$Ka = \frac{(H)\ (F)}{HF}$$

3. An increase in F^- ions will cause a shift in the equilibrium condition resulting in an increase in the numerator value for the equilibrium expression. In order to maintain Ka (ionization) constant for the acid involved) HF will increase due to a corresponding decrease in H^+ ions. Ka is maintained. Note that since H^+ decreases a pH increase also results.

4. Acids = proton donors

   Bases = proton acceptors

5. $NaOH + NCl \quad NaCl + H_2O$

6. a. T
   b. F

7. LeChateliers principle states that a reaction, at equilibrium, will adjust itself in such a way as to relieve any force, or stress, that disturbs the equilibrium.

8. Hydrogen ion

9.

	14	
	13	
	12	
	11	
	10	increases basic
	9	
	8	
pH	7	neutral
	6	
	5	
	4	increases acidic
	3	
	2	
	1	
	0	

10. 100 ml of NaOH are required.

11. H_2SO_4 - strong

H_2CO_3 - weak

HCl - strong

HNO_3 - strong

12. Buffers are substances in solution that offer a resistance to changes in pH as acids or bases are added to or formed within the solution.

13. Cl ion (chloride)

14. High concentrations may result from bacterial oxidation or organic matter with which the material has been in contact, and under the conditions of ground water, the carbon dioxide is not free to escape to the atmosphere.

15. Algae use carbon dioxide in their photosynthetic activity, and this removal is responsible for such high pH values.

16. True

17. $BaSO_4(s)$ $Ba^{++} + SO_4 =$

$K_{sp} = (Ba)\ (SO_4)$

K_{sp} = solubility product constant

18. If the sulfate (SO_4) concentration should be increased by addition from an external source the concentration of Ba ion must decrease and the amount of precipitated $BaSO_4$ must increase in order for K_{sp} to remain the same.

19. Adsorb

20. Osmosis is the movement of a solvent through a membrane which is impermeable to a solute. The direction of flow is from the more dilute to the more concentrated solution. Dialysis is the condition whereby ions are caused to pass through a wetted membrane of a particular permeability while large molecules of organic compounds or colloids are unable to pass through. Thus a separation of solutes by dialysis is achieved.

HANDOUT No. 1 - VALENCE

A. Common Radicals

Name	Formula	Valence
Hydroxyl	OH	-1
Nitrite	NO_2	-1
Nitrate	NO_3	-1
Sulfate	SO_4	-2
Carbonate	CO_3	-2
Bicarbonate	$H\ CO_3$	-1
Phosphate	PO_4	-3
Silicate	SiO_3	-2
Ammonium	NH_4	+1
Cyanide	CN	-1

B. Examples of compounds and ions of compounds

Calcium chloride $CaCl_2$ yields Ca (+2) and Cl (-1) ions

Calcium carbonate $CaCO_3$ yield Ca (+2) and CO_3 (-2) ions

Sodium chloride NaCl yields Na (+1) and Cl (-1) ions

Sodium nitrate $NaNO_3$ yield Na (+1) and NO_3 (-1) ions

Potassium chlorite K Cl yields K (+1) and Cl (-1) ions

Ammonium hydroxide NH_4OH yields NH_4 (+1) and OH (-1) ions

Magnesium sulfate $MgSO_4$ yields Mg (+2) and SO_4 (-2) ions

Calcium bicarbonate $Ca\ (HCO_3)_2$ yields Ca (+2) and HCO_3 (-1) ions

Sodium carbonate Na_2CO_3 yields Na (+1) and CO_3 (-2) ions

Calcium hydroxide $Ca\ (OH)_2$ yields Ca (+2) and OH (-1) ions

Ferric chloride $FeCl_3$ yields Fe (+3) and Cl (-1) ions

Cupric sulfate $CuSO_4$ yields Cu (+2) and SO_4 (-2) ions

Aluminum sulfate $Al_2(SO_4)_3$ yields Al (+3) and SO_4 (-2) ions

HANDOUT No. 2 - MOLECULAR WEIGHTS

Formula or molecular weight (M.W.) = the sum of all the atomic weights of the elements present in the formula of the compound.

Ex. NaCl atomic wt. of Na = 22.997

atomic wt. of Cl = 35.457

Molecular st. M.W. of NaCl = 58.454

Ex. $CaCO_3$ atomic wt. of Cu = 40.080

atomic wt. of C = 12.011

Molecular wt. M.W. of $CaCO_3$ = 100.088

Note that 3 atoms of oxygen are in the formula and for practical purposes 100.088 may be rounded off to 100.

Ex. $Ca(OH)_2$ atomic wt. of Ca = 40.080

atomic st. of O = 15.999 x 2 = 31.998

atomic wt. of H = 1.008 x 2 = 2.016

Molecular wt. M.W. of $Ca(OH)_2$ = 74.094

Ex. CH_4 atomic wt. of C = 12.011

atomic wt. of H = 1.008 x 4 = 4.032

Molecular wt. M.W. of CH_4 = 16.043

Mole - the quantity of a compound equal in weight to its molecules weight or formula weight; may contain fractions thereof.

Ex. 1 mole of $CaCO_3$ = 100.088 units, usually moles are expressed in gram amounts so 1 mole of $CaCO_3$ = 100.088 grams.

Likewise 1 mole of NaCl = 58.454 g.

1 mole of $Ca(OH)_2$ = 74.094 g.

.4 moles of $Ca(OH)_2$ = 74.094 g x .4 = 29.638 g.

.1 moles of $Ca(OH)_2$ = 74.094 g x .1 = 7.409 g.

2 moles of $Ca(OH)_2$ = 74.094 g. x 2 = 148.188 g.

Or given the quantity of a compound in grams find the number of moles it contains.

Ex. 38 grams of $Ca(OH)_2$ are used to make up a solution, how many moles would that be?

$$\frac{\text{Actual weight}}{\text{Molecular weight}} = \text{No. of moles } \frac{38 \text{ g.}}{74.094 \text{ g.}} = .513 \text{ moles}$$

Ex. 29 grams of NaOH (solid) are used to make up a solution, compute the number of moles.

Na atomic wt. = 22.997 g.

O atomic wt. = 15.999 g.

H atomic wt. = 1.008 g.

Molecular wt. NaOH = 40.004 g.

$$\text{Actual wt.} = \frac{29.0 \text{ g.}}{\text{M. W.}} \quad \frac{29.0 \text{ g.}}{40.004 \text{ g.}} = .725 \text{ moles of NaOH}$$

A molar solution of a compound is the number of moles per liter of solution (1M). A solution which contains a half mole of compound per liter of solution is designated as .5 M, and so on.

Ex. To make a 1 M (1molar) solution of NaOH in water 1 mole or 40.004 g. (solid) would be added to a container and water would be filled to make a total of 1 liter. For a .5M solution 40.004 x .5 or 20.002 g. would be used.

Ex. Prepare a .15 molar solution of Na_2CO_3 to find the number of grams of Na_2CO_3 required. .15 molar = .15 moles of Na_2CO_3.

M. W. of Na_2CO_3

Na at wt. = 22.997 x 2 = 45.994

C at wt. = 12.011

O at wt. = 15.99 x 2 = 31.998

M. W. Na_2CO_3 = 90.003

1 mole at Na_2CO_3 = 90.003 g.

.15 moles = 90.003 g. x .15 = 13.5 g.

Or 13.5 g. of Na_2CO_3 would be used per liter to prepare a .15 M solution.

HANDOUT NO. 3 - NORMALITY

Normality = number of equivalents of solute per liter of solution

$$\text{Normality (N)} = \frac{\text{No. of equivalents}}{\text{liter of solution}}$$

6N = 6 equivalents/liter

1N = 1 equivalent/liter

.16N = .16 equivalents/liter

$$\text{No. of equivalents} = \frac{\text{actual weight of the substance}}{\text{equivalent weight}}$$

$$\text{Equivalent weight} = \frac{\text{molecular weight}}{\text{valence}}$$

To determine equivalent weight, valence of the compound must be found. The product of the valence of the positive part of a compound times its subscript gives what is known as the number of replaceable hydrogens or net positive valence.

Example

$Ca^{++}CO_3$ = 2 x 1 = 2 valence

$Na^{+}\ Cl^{-}$ = 1 x 1 = 1 valence

$H^{+}_2\ SO_4$ = 1 x 2 = 2 valence

$Ca^{+}\ (HCO_3)^{-}$ = 2 x 1 = 2 valence

$H^{+}\ Cl^{-}$ = 1 x 1 = 1 valence

To determine equivalent weight of these compounds take the molecular weight/valence.

Example - $\frac{CaCO_3 \text{ molecular wt.} = 100}{\text{valence} = 2} = 50$ eq. wt.

If, for example 22.5 g. of $CaCO_3$ were dissolved to 1 liter of water its N is determined by:

Equivalent wt. of $CaCO_3$ = 50

No. of equivalents = $\frac{\text{actual wt. - -2.5}}{\text{eq. wt. - 50}}$ = .45

Normality = .45 equivalents/liter = .45 N

Normality and equivalents are useful since 1 equivalent of compound (A) will react exactly with 1 equivalent of compound (B), .3 equivalents with .3 equivalents and so on.

Reactant A + Reactant B = Product C + Product D

For example 1 equivalent of an acid will react with exactly 1 equivalent of a base.

1	2	3	4
HCl	+ NaOH	NaCl	+ H_2O
Acid	Base	Salt	Water

Using the basic equation: $N_1 \times V_1 = N_2 \times V_2$

Normality of reactant 1 x volume 1 (1) = normality of reactant 2 x volume 2 (1)

Suppose that 500 ml of a sample containing HCl is titrated with NaOH (.5 N) and 10 ml. of NaOH is required, find N of the HCl acid.

$N_1 \times V_1 = N_2 \times V_2$

$N_1 \times .5\ l = .5\ N \times .01\ l$

$N_1 = \frac{.5N \times .01\ l}{.51} = .01\ N\ HCl$